1. A chloride melts at 224˚C and reacts vigorously with water. State the type of bonding and structure present in this chloride and explain your reasoning.
   - Covalent AND simple/molecular
   - Low melting point/ reaction with water

2.

chloride	melting point/°C	difference between the electronegativities of the elements
NaCl	801	2.2
$SiCl_4$	–69	1.3

Explain the differences between the melting points of these two chlorides in terms of their structure and bonding. Refer to the difference between the electronegativities of the elements.

- NaCl structure and bonding: giant/lattice AND ionic
- $SiCl_4$ structure and bonding: simple/molecular AND covalent
- Explanation:
  - For NaCl, large difference in electronegativity of Na and Cl indicates electron transfer // ions
  - For SiCl4, smaller difference in electronegativity indicates sharing/ covalency with weak van der Waals' forces/ IM forces between molecules.

3.

$AlCl_3$

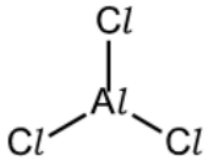
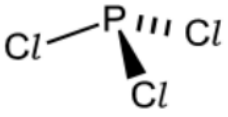
shape .....

angle .....

PCl_3

shape .....

angle .....

 trigonal planar 120°	 (trigonal) pyramid(al) 100–107°

4. Explain, in terms of structure and bonding, why the melting point of SiCl4 is much lower than that of NaCl.
    - SiCl4 simple/ molecular AND van der Waals'/ id–id/ London dispersion forces/ IMFs
    - NaCl ionic OR giant
    - Bonding in NaCl stronger than forces in SiCl4

5. Explain why the melting point of SiCl4 is higher than that of PCl3
    - SiCl4 has more electrons
    - Stronger van der Waals'/ id–id/ London dispersion forces/ IMFs

NOTE: for dot-and-cross diagrams, DO NOT draw the circles to represent shells. Write the symbol of each atom, surrounded by dots/crosses to represent electrons.

NOTE: for mp & bp, talk about the IMFs between molecules NOT strength of the covalent bonds!

6.

Copper, ice, silicon(IV) oxide, iodine and sodium chloride are all crystalline solids.

Complete the table with:

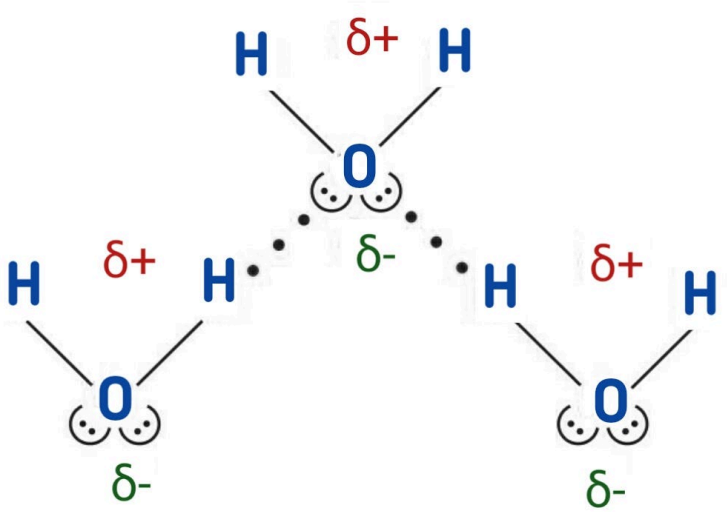
- the name of a type of bonding found in each crystalline solid,
- the type of lattice structure for each crystalline solid.

crystalline solid	type of bonding	type of lattice structure
copper		
ice		
silicon(IV) oxide		
iodine		
sodium chloride		

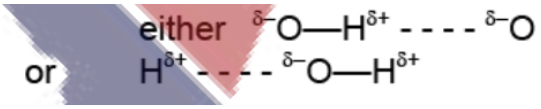
[5]

substance	type of bonding	type of lattice structure
copper	metallic	giant/metallic
ice	covalent OR hydrogen(-bonding) / H(–bonding)	hydrogen-bonded / simple / molecular
silicon(IV) oxide	covalent	giant (molecular) / macromolecular
iodine	covalent	simple / molecular
sodium chloride	ionic	giant / ionic

Diagram of hydrogen bonding in water:



- Mark 1: H-bond between O and H of different molecules
- Mark 2: minimum 3 partial charges in a row over 2 H2O molecules



- Mark 3: lone pair of electrons on O of H-bond, in line with H-bond.

7.

Some electronegativity values are shown.

element	electronegativity
aluminium	1.5
chlorine	3.0
iron	1.8

Use the data to suggest the nature of the bonding in iron(III) chloride. Explain your answer.

- Covalent
- small(er) difference in electronegativity between Fe and Cl (than between Al and Cl)

8. Name the force that has to be overcome in order to boil neon or argon and explain what causes it.
   - Van der Waals' forces/ London dispersion forces
   - Uneven electron distribution/ temporary dipole
   - Induced dipole-dipole attraction

9. Explain why argon has a higher boiling point than neon
   - More electrons
   - More polarisable/ greater attraction/ stronger IMFs

10. Define the term *electronegativity*
    - power of an atom to attract
    - a bonding pair of electrons in a covalent bond towards itself

11. Name and describe the type of bonding expected to be found between particles with equal electronegativities.
    - Covalent
    - Shared <u>pair</u>(s) of electrons OR overlap of orbitals

    OR
    - Metallic
    - Positive ions/cations surrounded by delocalised electrons

12. Name and describe the type of bonding expected to be found between particles with very different electronegativities.
    - ionic/electrovalent
    - Electrostatic attraction between oppositely charged (+ve and -ve) ions

13. Explain why the boiling points of fluorine and argon are similar
    - Similar strength/amount/number of intermolecular forces/ van der Waals' forces/ induced dipole/ London dispersion forces.
    - Therefore similar energy needed.

14. Explain why boiling point of hydrogen chloride is higher than that of fluorine.
    - HCl is polar/ has a dipole AND F2 is non-polar/ has no dipole OR
      HCl has permanent dipole-dipole attractions between molecules AND F2 has induced dipole-dipole attractions between molecules.
    - Thus more energy needed for HCl than for F2 OR
      pd-pd forces stronger than id-id forces OR
      IMFs/ vdWs in HCl stronger than in F2

15. Explain why methanol has highest boiling point (in comparison to other given molecules)
    - Hydrogen bonding between methanol molecules.
    - Stronger than the IMFs/ vdWs in the others // is the strongest intermolecular force.

16. Explain why O has higher Pauling electronegativity value than S
    - O lower nuclear charge / lower proton number
    - O has (one) fewer shell than S / less shielding
    - greater attraction (for nucleus) in O

17. Suggest the shape of a molecule of H_2S
    Non-linear

18.

(e) (i) Complete Table 1.2 by placing a tick (✓) to show which of the compounds have molecules with an overall dipole moment.

Table 1.2

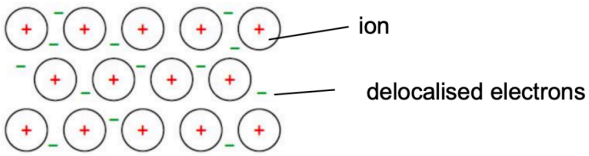
compound	O=C=O	O=S=O	S=C=S	S=C=O
overall dipole moment				

[2]

compound	O=C=O	O=S=O	S=C=S	S=C=O
overall dipole moment		✓		✓

19.

Draw a labelled diagram to show the bonding and structure of the Group 2 metals at room temperature.



M1 diagram showing minimum of 4 particles (in total in 2 rows)

- circles containing X^{n+} do not have to be labelled
-  must be labelled as 'ion' OR empty circles / circles with X must be labelled + ion / positive ion / cation / X^{n+}

AND
with the circles surrounded by electrons shown as e^- / – **OR** little circles labelled electrons

---

M2 label / legend showing **delocalised** electrons

20. Explain why Mg has a higher electrical conductivity than Na
Mg has more delocalised e– (than Na).

21.

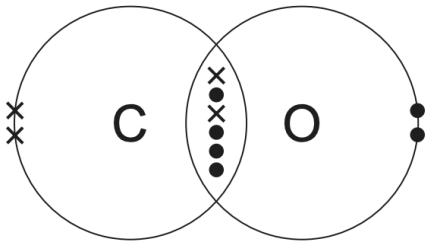
Identify all types of van der Waals' forces that are present between molecules of CH_3CH_2CHO.

instantaneous dipole—induced dipole / id—id
AND permanent dipole—permanent dipole / pd—pd

22. P_4O_6 is a white solid that has a melting point of 24°C. Solid P_4O_6 reacts with water to form H_3PO_3. Deduce the type of structure and bonding shown by P_4O_6. Explain your answer.
    - Structure = simple molecular because it has low melting point
    - Bonding = covalent because it is hydrolysed

23.

The diagram shows the arrangement of outer electrons in a molecule of CO(g).



State one similarity and one difference in the way the atoms in a carbon monoxide molecule are bonded together compared to the atoms in a nitrogen molecule.

- both make triple (covalent) bond / 3 shared pairs of electrons
- one bond in CO is coordinate / dative covalent / formed by donating a pair of electrons from O (to C)

24.

	C	N	O
electronegativity	2.5	3.0	3.5

Use the electronegativity values and relevant details from the *Data Booklet* to complete the table below.

	N_2	CO
number of electrons per molecule		
type(s) of intermolecular (van der Waals') force		

	N_2	CO
number of electrons per molecule	14	14
type of van der Waals'	temporary / instantaneous dipole–induced dipole	permanent dipoles–(permanent) dipoles (and temporary / induced / instantaneous dipoles)

25. Suggest why CO(g) is more reactive than N2(g).
    CO / it is a polar molecule / it has a (permanent) dipole (but N2 is non-polar).

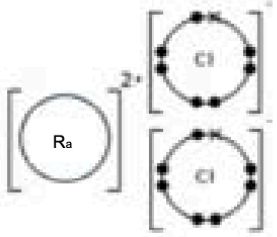
1. Describe how triple covalent bond forms in a N2(g) molecule. Refer to orbital overlap and hybridisation in your answer.
   - 1 σ bond and 2 π bonds
   - sp hybridisation in each N atom
   - σ bonds form direct/end-on overlap of orbitals AND π bonds form sideways/lateral overlap of orbitals

2. Mg has mp of 650˚C and high electrical conductivity. Explain these properties of Mg by referring to its structure and bonding.
   High mp
   - many strong metallic bonds OR
   - many strong (electrostatic) attractions between cations and delocalised electrons OR
   - strong bonds in giant metallic structure

   Electrical conductivity
   - delocalised electrons can move through the structure

3. Draw a dot-and-cross diagram to show the arrangement of outer electrons in $RaCl_2$.

   Ra^{2+} and 2 x Cl^- 0 electrons surrounding 2Ra **AND** 8 electrons surrounding Cl

   NOTE: the outermost shell of the metal should have 0 electrons!!!

4. Tennessine, Ts, is an unstable man-made element. It is found below astatine, At, in Group 17. The chemical properties of Ts and its compounds have only been predicted.
   Some scientists predict that Ts has properties typical of metals like copper.

Complete Table 2.2 with:

- the predicted melting point of tennessine
- the lattice structure of solid chlorine, bromine and tennessine.

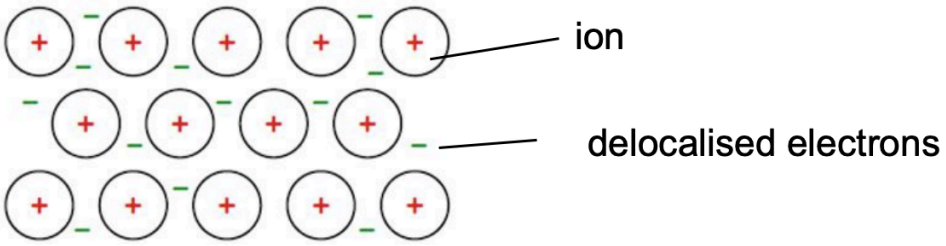
Assume that Ts has properties typical of metals like copper.

Table 2.2

element	chlorine	bromine	tennessine
melting point / °C	–101	–7.2	
lattice structure of crystalline solid			

element	chlorine	bromine	tennessine
melting point / °C	–101	–7.2	> 250
lattice structure of crystalline solid	simple molecular	simple molecular	giant metallic

5. Draw a labelled diagram to show the bonding in magnesium metal.



- minimum of 4 particles in rows – circles
- circles must be labelled as 'ion' OR empty circles / circles with Mg must be labelled + ion / positive ion / cation / Mg^{n+}
- circles surrounded by electrons shown as e– / –
- label showing <u>delocalised</u> electrons

6. SCl2 is a cherry-red liquid that reacts vigorously with water to form an acidic solution. Use this information to deduce the bonding and structure shown by SCl2. Explain your answer.
   - simple / molecular structure <u>because</u> it has a low melting / boiling point covalent bonding <u>because</u> it is hydrolysed

7.

Fig. 3.1 shows the two structural isomers of S_2Cl_2.

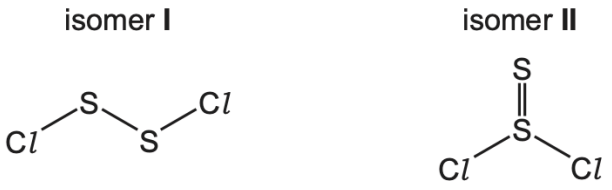
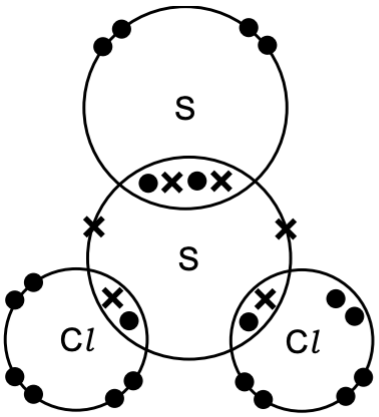


Fig. 3.1

Suggest a value for the Cl–S–S bond angle in isomer I. Explain your answer.

- 90° < Cl—S—S ⩽ 108°
- sulfur has two lone pairs (of e–) (and two bonding pairs) AND repulsion from lone pairs (greater)

8. Draw a dot-and-cross diagram to show the bonding in isomer II. Show outer shell electrons only.



9. 

The above organic compound has melting point –94°C and boiling point +49°C. Explain these properties by referring to the type of van der Waals' forces between molecules.

- (m.& b.pts are low because) weak intermolecular forces / weak van der Waals
- (only) London / dispersion forces / instantaneous dipole-induced dipole

10. The melting point of silicon is 1410°C. The melting point of sulfur is 113°C. Explain this difference.

- Si giant covalent AND S is simple molecular
- Si involves breaking covalent bonds AND S involves breaking IMFs

- more energy needed to break covalent bonds than IMFs OR covalent bonds are strong AND IMFs are weak

11. Define free radical
    - species with 1 or more unpaired electrons

12.

(e) (i) Complete Table 1.2 by placing a tick (✓) to show which of the compounds have molecules with an overall dipole moment.

Table 1.2

compound	O=C=O	O=S=O	S=C=S	S=C=O
overall dipole moment				

compound	O=C=O	O=S=O	S=C=S	S=C=O
overall dipole moment		✓		✓

13. Explain why Mg has a higher electrical conductivity than Na.
    - Mg has more delocalised e– (than Na)

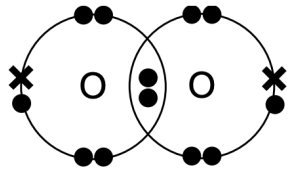
14. Identify all types of van der Waals' forces that are present between molecules of CH_3CH_2CHO
    - Instantaneous dipole–induced dipole AND permanent dipole–permanent dipole

15. Silicon reacts with nitrogen gas to form Si_3N_4. Si_3N_4 is a solid with a melting point of 1900°C. It is insoluble in water and does not conduct electricity when molten. Suggest the type of bonding in and structure of Si_3N_4. Explain your answer.
    - giant molecule
    - strong covalent bonds between atoms / particles
    - no mobile charged particles / carriers

16.

The peroxide ion, O_2^{2-}, has a single covalent bond between the two oxygen atoms. Each oxygen atom carries a negative charge.

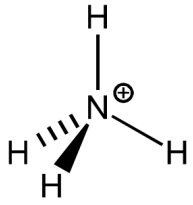
Draw a 'dot-and-cross' diagram for the peroxide ion. Show outer electrons only.



M1: bonding pair between the two O

M2: total of 14 electrons distributed equally between the two O

17. Draw the three-dimensional shape of the ammonium ion, NH_4^+ . Give the bond angle.



AND $109\frac{1}{2}(^\circ)$

18. Explain why TeFx is octahedral with bond angles of 90°.
- TeFx has 6 bonding pairs of electrons, and 0 lone pairs OR coordination number = 6
- Bonding pairs repel equally

19.

The melting points of some solids are shown in Table 1.1.

Table 1.1

solid	melting point/K
magnesium	923
phosphorus	317
sodium chloride	1074
sulfur	392

a. Explain the difference in the melting points of magnesium and sodium chloride. [1 mark]
- bonds in NaCl are stronger than bonds in Mg

b. Explain the difference in the melting points of phosphorus and sulfur in terms of structure and bonding.
- S8 / molecules of sulfur have more electrons than P4 / molecules of phosphorus

- Sulfur has stronger instantaneous dipole –induced dipole forces than phosphorus

20.

The melting points of ice and ammonia are shown in Table 1.2.

Table 1.2

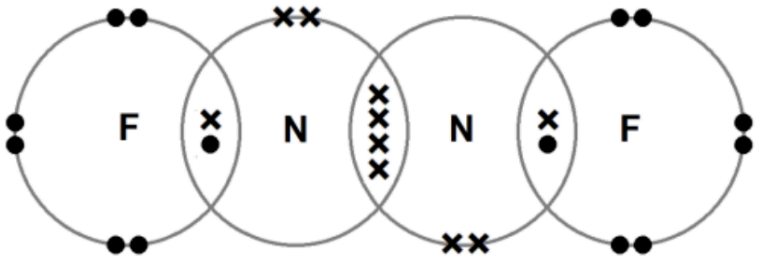
solid	melting point/K
ice	273
ammonia	195

Suggest **two** reasons for the difference in the melting points of ice and ammonia.

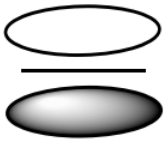
- O is more electronegative than N
- Ice has 2 hydrogen bonds per molecule while ammonia has 1 hydrogen bond per molecule

21. Bi2O3 is a yellow insoluble solid that melts at 1090K. The molten compound conducts electricity. Deduce the structure and bonding of Bi2O3. Explain your answer.
    - giant AND has a high melting point
    - ionic AND conducts electricity (only) when molten

22. Bonding in N2F2:



Draw a diagram of the π bond between the N atoms in N2F2 and describe how it forms.



- overlap of two 2p orbitals side-on / above and below the plane

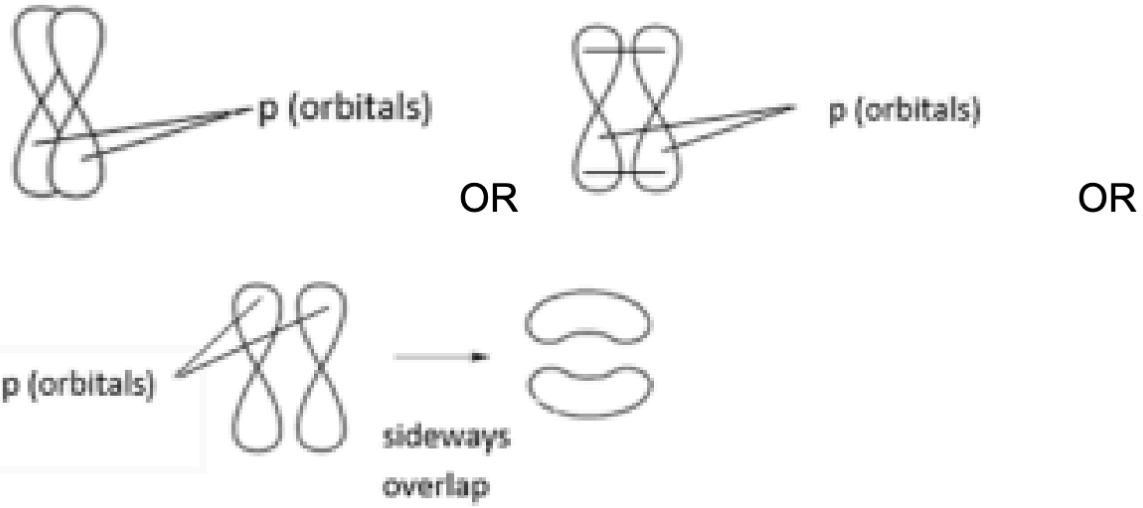
23. Define covalent bond

- Electrostatic attraction between nuclei of two atoms and shared pair of electrons

24. Draw labelled diagrams to show, in terms of orbital overlap, how the σ and π bonds are made in a C=C bond.

<u>σ bond</u>

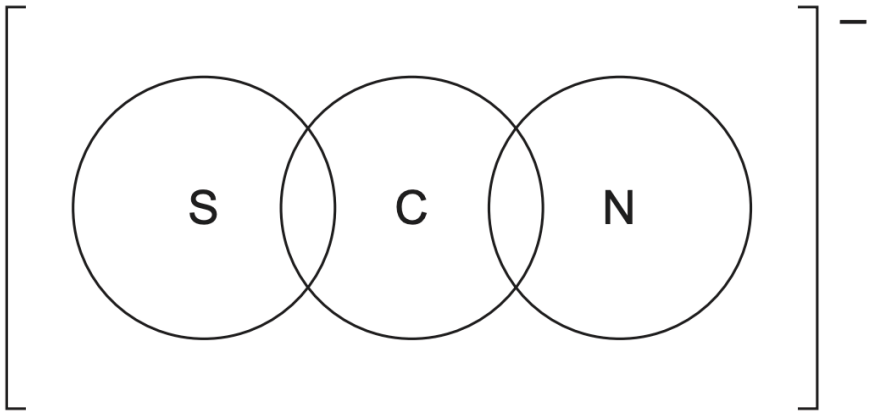
<u>π bond</u>

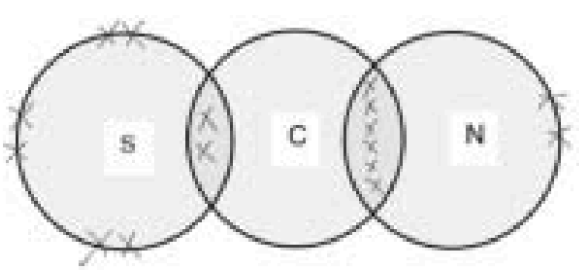


25.

$SCN^-(aq)$ is colourless.

Complete the dot-and-cross diagram in Fig. 3.1 to show the arrangement of outer electrons in an SCN^- ion.

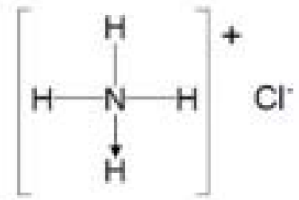




M1 6 electrons between C and N **AND** 2 electrons between C and S
M2 3 lone pairs around S **AND** none around C **AND** 1 lone pair around N

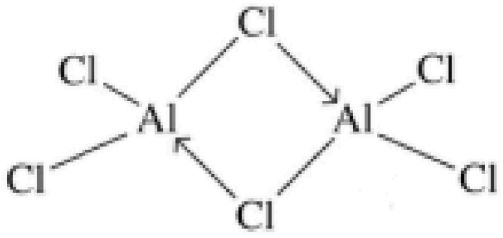
NOTE: when you have a CN bond, assume that it is a triple bond!

26. Draw a diagram to show the ionic, covalent and coordinate bonding present in a formula unit of NH4Cl.



- ionic bond between NH4+ and Cl – AND 4 N—H covalent bonds OR 4 shared pairs of electrons in NH4
- 1 N—H coordinate bond OR arrow on bond/line in correct direction

27. Draw a diagram that clearly shows all the types of bond present in Al2Cl6(g).



28.

0.020 mol of element **Z** reacts with excess Cl_2 to form 0.020 mol of a liquid chloride.

The liquid chloride has formula **Z**Cl_n, where *n* is an integer.

ZCl_n reacts vigorously with water at room temperature to give an acidic solution and a white solid.

When excess $AgNO_3$(aq) is added to the solution, 11.54 g of AgCl(s) forms.

Suggest the type of bonding and structure shown by ZCln.

- simple/molecular AND covalent

29. CS2 is a liquid under room conditions, while CO2 is a gas. Explain what causes the difference in the physical properties between CS2 and CO2.
    - CS2 has more electrons
    - Thus stronger induced dipole forces between molecules

30. Describe how sp hybridised orbitals are formed.
    - mixing / combination of one s and one p orbital

31. Sketch a diagram to show how two sp hybrid orbitals can form a sigma (σ) bond

M1

M2

32.

Phosphorus is a reactive Period 3 element.

(a) Phosphorus has several allotropes. Details of two allotropes are given.

allotrope of phosphorus	formula	melting point / °C
white	P_4	44
red	P	590

White phosphorus and red phosphorus both have covalent bonding.
Suggest the types of structure shown by white phosphorus (P4) and red phosphorus (P). Explain why red phosphorus (P) has a higher melting point than white phosphorus (P4).

- Structure of P4: simple molecular
- Structure of P: giant molecular
- Explanation: weak IMFs overcome in P4 & strong covalent bonds broken in P

NOTE: when asked for structure, say simple molecular / giant molecular
When asked for structure AND bonding, say simple covalent / giant covalent

33. Identify the lattice structure shown by graphite.
    - Giant molecular

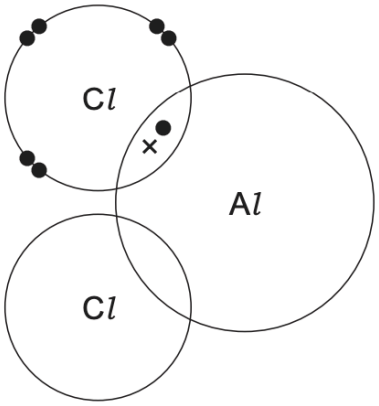
34. SiO2 is a white solid that melts above 1700°C. SiCl4 is a colourless liquid at room temperature. Explain the difference in the melting points of these two compounds with reference to their structure and bonding.
    - SiO2 is giant covalent AND SiCl4 is simple covalent / simple molecular
    - so need less energy to overcome IMF's in SiCl4, compared to the bonds in SiO2

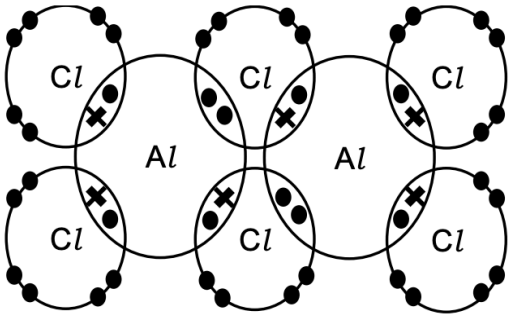
35. POCl 3 has a melting point of 1°C and a boiling point of 106°C. POCl3 reacts vigorously with water, forming misty fumes and an acidic solution. Explain how this information suggests the structure and bonding of POCl3 is simple covalent.
    - relatively low melting and boiling points suggests weak intermolecular / VdW forces
    - vigorous reaction with water suggests hydrolysis

36. Identify the species that are electrostatically attracted to one another in:
    - An ionic bond: oppositely charged ions / cations AND anions
    - A covalent bond: shared pair(s) of electrons AND two nuclei

37.

Complete the dot-and-cross diagram to show the bonding in Al_2Cl_6.



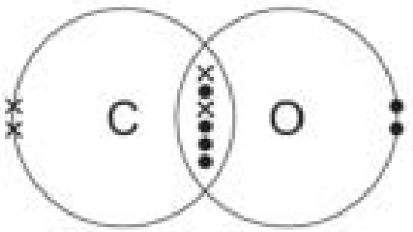


38. SCl_2 has Mr = 103.1 and is a liquid at room temperature. SBr_2 has Mr = 191.9 and is a gas at room temperature. Explain the difference in the physical state of SCl_2 and SBr_2. Give your answer in terms of intermolecular forces.
    - sum of instantaneous dipole–induced dipole attractions and permanent dipole attractions in SCl_2 is greater than those in SBr_2 OR permanent dipole-permanent dipole (forces) in SCl_2 are stronger OR intermolecular forces are stronger in SCl_2
    - SCl_2 molecule / bond is more polar than SBr_2 OR Cl is more electronegative than Br so SCl_2 has stronger permanent dipoles

39. Explain, in terms of the structure of the crystalline solid, why lithium sulfide has a high melting point.
    - giant
    - many strong forces of attraction between oppositely charged ions OR many strong ionic bonds

40. Explain what is meant by coordinate bond.
    - covalent bond with both electrons provided from the same / one species OR
    - shared pair of electrons are provided from the same species / one atom

41. Draw a 'dot-and-cross' diagram to show the arrangement of outer electrons in CO. Show the electrons belonging to the C atom as ×. Show the electrons belonging to the O atom as ●.



<u>NOTE: comparing boiling points of different compounds</u>

<u>H_2O and HF</u>

- Bp of H_2O and HF are relatively high since they have hydrogen bonds.
- Bp of H_2O > HF
- Each hydrogen bond in HF is stronger than each hydrogen bond in HF because F is more electronegative than O.
- However, H_2O is more extensively hydrogen bonded than HF. H_2O has an average of 2 hydrogen bonds per molecule (2 lp on O & 2 H atoms) while HF has an average of 1 hydrogen bond per molecule (3 lp on O but only 1 H atom)

<u>HCl and PH_3</u>

- They have similar bp.
- They have same number of electrons = strength of van der Waals' forces / id-id forces is similar = similar amount of energy needed to break them.

<u>HCl and H2S</u>

- Bp of H2S > HCl
- H2S has more contact points/ larger surface area = greater id-id attraction = larger boiling point